Chemistry Quiz
Thermodynamics
Instructor's Version B

Answer Sheet

1.	D
2.	C
3.	D
4.	B
5.	E
6.	A
7.	D
8.	C
9.	C
10.	E

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Free-Response Questions

Time Limit: You have 30 minutes to complete this section.
Scoring: 60 points total; Question 1 is worth 40 points and Question 2 is worth 20 points; points for each section are indicated within the questions.
Calculator Use: You may use a calculator on both questions.

1. For the reaction: $2CO\ (g) + O_2\ (g) \rightarrow 2CO_2\ (g)$

A. **What would you predict as the sign for ΔS? Explain. (5 points)**

ΔS is negative; 3 moles of gaseous reactant → 2 moles of gaseous product

B. Using the thermodynamic data provided below, determine the free energy change of the reaction. (10 points)

ΔG_f° CO(g) = – 137.3 kJ/mol
ΔG_f° O_2(g) = 0 kJ/mol
ΔG_f° CO_2(g) = – 394.4 kJ/mol

ΔG°_{rxn} = 2(-394.4) – 2(-137.3) = – 514.2 kJ

C. Is the reaction spontaneous? Explain. (5 points)

Yes, delta G is negative.

D. Do you expect this reaction to be endothermic or exothermic? Explain. (10 points)

This reaction should be exothermic. Since delta G is negative, but delta S is negative, then the delta H factor must be negative.

2. For the reaction: $CO\ (g) + H_2O\ (g) \Leftrightarrow CO_2\ (g) + H_2\ (g)$
ΔH° = – 41.2 kJ ΔS° = – 42.0 J/K

A. Calculate ΔG° at 25°C. (10 points)

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ = – 41,200 J – (298K)(– 42.0 J/K) = – 28,700 J = –28.7 kJ

B. Determine the temperature at which this reaction will change from being spontaneous to non-spontaneous. (5 points)

$T = \Delta H^\circ / \Delta S^\circ$ = 41,200 / 42.0 = 981 K or 708 °C

C. Why is this temperature an approximation? (2 point)

ΔH° and ΔS° are temperature-dependent. Using these at temperatures other than 25 °C is an approximation.

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D. Given the starting partial pressures shown below, determine whether the reaction is at equilibrium (at 25°C), and if not, indicate in which direction it will shift. Show all work clearly to back up your answers. (13 points)
P_{CO} = 1 atm P_{H2O} = 0.5 atm P_{CO2} = 0.1 atm P_{H2} = 0.1 atm

$\Delta G^\circ = - RT\ln K$

$\ln K = \Delta G^\circ / -RT = (-28{,}700\ J) / - (8.314\ J/mol\ K)(298) = 11.58$

$K_c = K_p$ (since the moles of gas product = moles of gas reactant)

So $K_p = e^{11.58} = 1.07 \times 10^5$

$Q = (0.1)(0.1) / (1)(0.5) = 0.02$

$Q << K$, so the system shifts to the right to make more product to reach equilibrium.

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